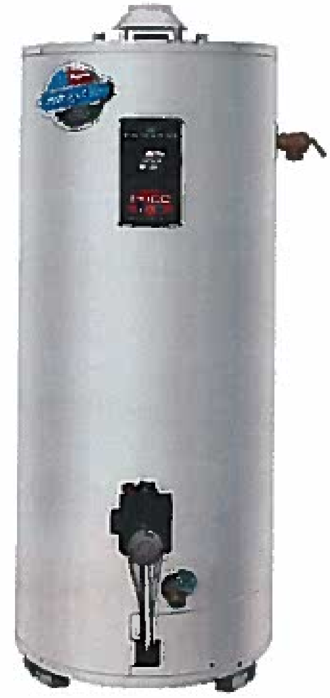
Question 1

An average hot water tank will hold 300.0 L of water needs to be kept at 65.0°C in order to kill any bacteria in the water. If you are filling a new tank with water at a temperature of 12.5°C, what mass of natural gas (methane) must be burned in a complete combustion reaction in order bring your tank up to code?



$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2\,H_2O_{(g)}$

Solution 1

(1) $\Delta H^\circ_x = \sum n\Delta H^\circ_{f(p)} - \sum n\Delta H^\circ_{f(r)}$

$= [(2)(-241.8) + (1)(-393.5)] - [(1)(-74.6) + 2(0)]$

$= (-877.1) - (-74.6)$

$= -802.5$ kJ/mol of CH_4

(2) $q_{water} = mc\Delta T$

$= (300\,000)(4.18)(52.5)$

$= 65\,835\,000$ J

or

65835 kJ

(3) $\Delta H = n\Delta H_x$

$n = \dfrac{\Delta H}{\Delta H_x}$

$= \dfrac{-65835}{-802.5}$

$= 82.04\ldots$ mol

(4) $m = nM$

$= (82.23)(16.05)$

$= 1316.7$ g

$\therefore \Delta H = -65835$ kJ

Question 2

The Thermite reaction can be used to produce molten iron for welding railway tracks together.

$$Fe_2O_{3(s)} + 2\,Al_{(s)} \rightarrow Al_2O_{3(s)} + 2\,Fe_{\cancel{(l)}\,(s)}$$

Calculate the amount of heat produced using 45.0 g of aluminum and excess iron (III) oxide.

Solution 2

① $\Delta H^\circ_x = [2(0) + (1)(-1675.7)] - [2(0) + (1)(-824.2)]$

$= -851.5 \text{ kJ} \rightarrow \Delta H^\circ_x = -425.75 \text{ kJ/mol}$

② $n_{Al} = \frac{m}{M}$

$= \frac{45.0}{26.98}$

$= 1.6679 \text{ mol}$

③ $\Delta H = n\Delta H_x$

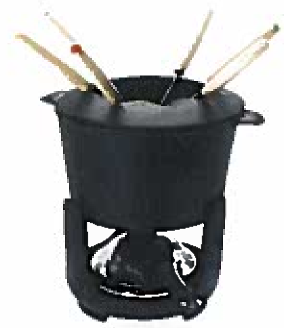
$= (1.6679)(-425.75)$

$= -710.109 \text{ kJ}$

$= -7.10 \times 10^2 \text{ kJ}$

Question 3

It's 1976 and you are throwing a fondue party! You need to heat 800.0 mL of vegetable oil to 375°F (190.0°C) using an ethanol burner. Assuming a perfect transfer of heat from the burner into the oil, how much ethanol is required. Assume your kitchen is at SATP and the specific heat capacity of vegetable oil is 2.000 J/g•°C



Density of veg oil $= 0.92\ g/mL$

$$C_2H_5OH_{(l)} + 3O_2 \xrightarrow{\Delta} 2CO_{2(g)} + 3H_2O_{(g)}$$

Solution 3

$(800.0\ mL)(0.92\ g/mL) = 736\ g$

① $q = mc\Delta T$
$= (736)(2.000)(165)$
$= 242\ 880\ J$

② $\therefore \Delta H = -242\ 880\ J$

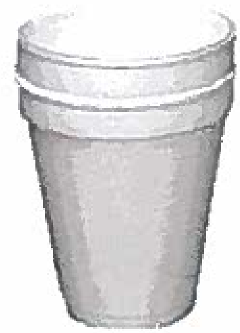
③ $\Delta H_f^\circ = \Sigma n\Delta H_f^\circ(\text{prod}) - \Sigma n\Delta H_f^\circ(\text{react})$
$= [(2)(-393.5) + (3)(-241.8)] - [(1)(-277.6) + 3(0)]$
$= (-1512.4) - (-277.6)$
$= -1234.8\ kJ/mol$

④ $\Delta H = n\Delta H_f$
$n = \dfrac{\Delta H}{\Delta H_f}$
$= \dfrac{-242.88\ kJ}{-1234.8\ kJ}$
$= 0.197\ mol$

⑤ $m = nM$
$= (0.197)(46.0$
$= 9.08\ g$

Question 4

100.0 mL of 3.0 M HCl is neutralized with can be neutralized with exactly 200.0 mL of 1.5 M NaOH. If this reaction is done in a coffee cup calorimeter, what temperature change would you expect to observe?



Solution 4

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

0.30 mol   0.30 mol   0.30 mol   0.30 mol

$$\Delta H = \sum n \Delta H_f^\circ \text{(prod)} - \sum n \Delta H_f^\circ \text{(react)}$$

$$= [(0.30)(-285.8) + (0.30)(-411.2)] - [(0.30)(-92.3) + (0.30)(-425.6)]$$

$$= [(-85.74) + (-123.36)] - [(-27.69) + (-127.68)]$$

$$= (-209.1) - (-155.37)$$

$$= -53.73 \text{ kJ}$$

$\therefore q = +53.73$ kJ $q = mc\Delta T$

$= 53730$ J $53730 = (300.0)(4.19)\Delta T$

$\Delta T = 42.7\ ^\circ C$

Question 5

A lab technician places a 5.00 g food sample into a bomb calorimeter with a heat capacity of 9.23 kJ/°C. The initial temperature of the calorimeter system is 21.0°C. After burning the food, the final temperature of the system is 32.0° C. How much thermal energy is released by the combustion of the food in kilojoules per gram?

Follow-up question!

Calories on food packages are actually kilocalories (kcal).

If 1 kj = 0.2388 kcal, what is the caloric value of the food sample?

Solution 5

$q = mc\Delta T$

$q = C\Delta T$

$= (9.23\ \frac{kJ}{°C})(11.0\ °C)$

$= 101.53\ kJ$

$\frac{101.53}{5} = 20.3\ kJ/g$

$20.3\ \frac{kJ}{g} \times \frac{0.2388\ kcal}{1\ kJ}$

$= 4.85\ kcal/g$

$4.85\ \frac{kcal}{g} \times 45\ g$

$= 218.25\ kcal$